Chapter 3 - Atomic Structure and the Periodic Table

Inside atoms

Atomic number and mass number
Ions
Isotopes
The mass spectrometer
Relative atomic mass

The periodic table

Electron configuration and the periodic table
Group 1 - the alkali metals
Group 2 - the alkaline earth metals
Flame colours
Group 7 - the halogens
Group 0 - the noble gases
Transition elements
The position of hydrogen

The arrangement of electrons in atoms

Checklist

Additional questions

We have already seen in Chapter 2 that everything you see around you is made out of tiny particles called atoms (Figure 3.1). When John Dalton developed his atomic theory, about 200 years ago (1807/1808), he stated that the atoms of any one element were identical and that each atom was `indivisible'. Scientists in those days believed that atoms were solid particles like marbles.

However, in the last hundred years or so it has been proved by great scientists, such as Niels Bohr, Albert Einstein, Henry Moseley, Joseph Thomson, Ernest Rutherford and James Chadwick, that atoms are in fact made up of even smaller `sub-atomic' particles. The most important of these are **electrons**, **protons** and **neutrons**, although 70 sub-atomic particles have now been discovered.

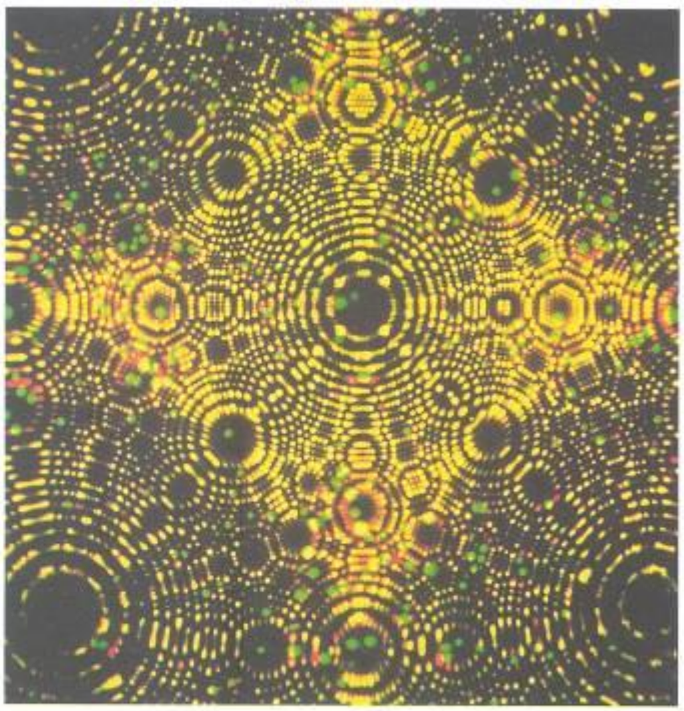


Figure 3.1 Atoms. A field ion micrograph of atoms of the metal element iridium. The tiny dots are the locations of individual atoms, and the ring-like patterns are facets of a single crystal.

Inside atoms

The three sub-atomic particles are found in distinct and separate regions. The protons and neutrons are found in the centre of the atom, which is called the nucleus. The neutrons have no charge and protons are positively charged. The nucleus occupies only a very small volume of the atom but is very dense.

The rest of the atom surrounding the nucleus is where electrons are most likely to be found. The electrons are negatively charged and move around very quickly in **electron shells** or **energy levels**. The electrons are held within the atom by an **electrostatic force of attraction** between themselves and the positive charge of the protons in the nucleus (Figure 3.2).

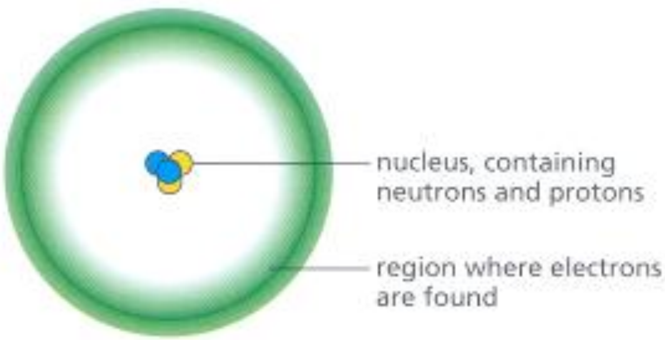


Figure 3.2 Diagram of an atom.

About 1837 electrons are equal in mass to the mass of one proton or one neutron. A summary of each type of particle, its mass and relative charge is shown in Table 3.1. You will notice that the masses of all these particles are measured in **atomic mass units (amu)**. This is because they are so light that their masses cannot be measured usefully in grams.

Table 3.1 Characteristics of a proton, a neutron and an electron.

Particle	Symbol	Relative mass/amu	Relative charge
Proton	p	1	+1
Neutron	n	1	0
Electron	e	1/1837	-1

Although atoms contain electrically charged particles, the atoms

themselves are electrically neutral (they have no overall electric charge). This is because atoms contain equal numbers of electrons and protons. For example, the diagram in Figure 3.3 represents the atom of the non-metallic element helium. The atom of helium possesses two protons, two neutrons and two electrons. The electrical charge of the protons in the nucleus is, therefore, balanced by the opposite charge of the two electrons.

Atomic number and mass number

The number of protons in the nucleus of an atom is called the **atomic number** (or proton number) and is given the symbol **Z**. Hence in the diagram shown in Figure 3.3, the helium atom has an atomic number of 2, since it has two protons in its nucleus. Each element has its own atomic number and no two different elements have the same atomic number. For example, a different element, lithium, has an atomic number of 3, since it has three protons in its nucleus.

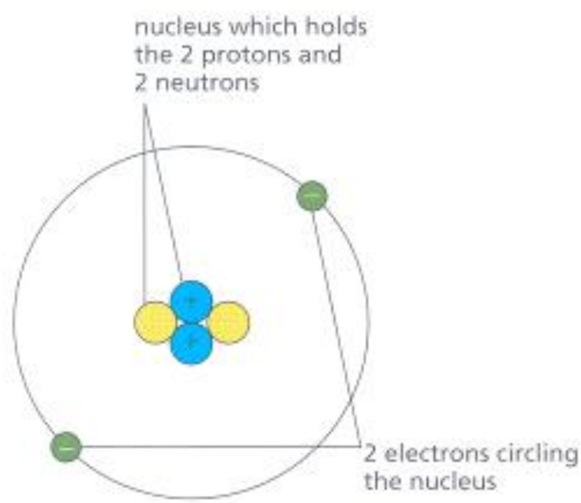


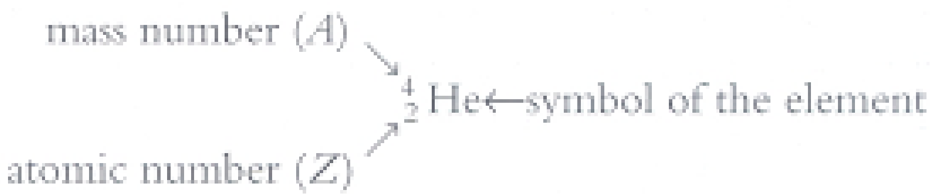
Figure 3.3 An atom of helium has two protons, two electrons and two neutrons.

Neutrons and protons have a similar mass. Electrons possess very little mass. So the mass of any atom depends on the number of neutrons and protons in its nucleus. The total number of protons and neutrons found in the nucleus of an atom is called the **mass number** (or nucleon number) and is given the symbol **A**.

mass number (A) = atomic number (Z) + number of neutrons

Hence, in the example shown in Figure 3.3 the helium atom has a mass number of 4, since it has two protons and two neutrons in its nucleus. If we consider the metallic element lithium, it has three protons and four neutrons in its nucleus. It therefore has a mass number of 7.

The atomic number and mass number of an element are usually written in the following shorthand way:



The number of neutrons present can be calculated by rearranging the relationship between the atomic number, mass number and number of neutrons to give:

number of neutrons = mass number (A) - atomic number (Z)

For example, the number of neutrons in one atom of $^{24}_{12}Mg$ is:

24 – 12 = 12
(A) (Z)

and the number of neutrons in one atom of $^{207}_{82}Pb$ is:

207 – 82 = 125
(A) (Z)

Table 3.2 shows the number of protons, neutrons and electrons in the atoms of some common elements.

Table 3.2 Number of protons, neutrons and electrons in some elements.

Element	Symbol	Atomic number	Number of electrons	Number of protons	Number of neutrons	Mass number
Hydrogen	H	1	1	1	0	1

Helium	He	2	2	2	2	4
Carbon	C	6	6	6	6	12
Nitrogen	N	7	7	7	7	14
Oxygen	O	8	8	8	8	16
Fluorine	F	9	9	9	10	19
Neon	Ne	10	10	10	10	20
Sodium	Na	11	11	11	12	23
Magnesium	Mg	12	12	12	12	24
Sulphur	S	16	16	16	16	32
Potassium	K	19	19	19	20	39
Calcium	Ca	20	20	20	20	40
Iron	Fe	26	26	26	30	56
Zinc	Zn	30	30	30	35	65

Ions

An ion is an electrically charged particle. When an atom loses one or more electrons it becomes a positively charged ion. For example, during the chemical reactions of potassium, each atom loses an electron to form a positive ion, K^+.

$$\begin{array}{rr} & 19\ \text{protons} = 19+ \\ {}_{19}K^+ & \underline{18\ \text{electrons} = 18-} \\ & \text{Overall charge} = 1+ \end{array}$$

When an atom gains one or more electrons it becomes a negatively charged ion. For example, during some of the chemical reactions of chlorine it gains an electron to form a negative ion, Cl^-.

$$\begin{array}{rr} & 17\ \text{protons} = 17+ \\ {}_{17}Cl^- & \underline{18\ \text{electrons} = 18-} \\ & \text{Overall charge} = 1- \end{array}$$

Table 3.3 shows some common ions. You will notice from Table 3.3 that:

- some ions contain more than one type of atom, for example NO_3^-

- an ion may possess more than one unit of charge (either negative or positive), for example Al^{3+}, O^{2-} or SO_4^{2-} .

Table 3.3 Some common ions.

Name	Formula
Lithium ion	Li^{+}
Sodium ion	Na^{+}
Potassium ion	K
Magnesium ion	Mg^{2-}
Calcium ion	Ca^{2+}
Aluminium ion	Al^{3+}
Zinc ion	Zn^{2-}
Ammonium ion	NH_4^{+}
Fluoride ion	F^{-}
Chloride ion	Cl^{-}
Bromide ion	Br^{-}
Hydroxide ion	OH^{-}
Oxide ion	O^{2-}
Sulphide ion	S^{2-}
Carbonate ion	CO_3^{2-}
Nitrate ion	NO_3^{-}
Sulphate ion	SO_4^{2-}

Isotopes

Not all of the atoms in a sample of chlorine, for ex-ample, will be identical. Some atoms of the same element can contain different numbers of neutrons. Atoms of the same element which have different numbers of neutrons are called **isotopes**.

The two isotopes of chlorine are shown in Figure 3.4. Generally, isotopes behave in the same way during chemical reactions. The only effect of the extra neutrons is to alter the mass of the atom and properties which depend on it, such as density. Some other examples of atoms with isotopes are shown in Table 3.4.

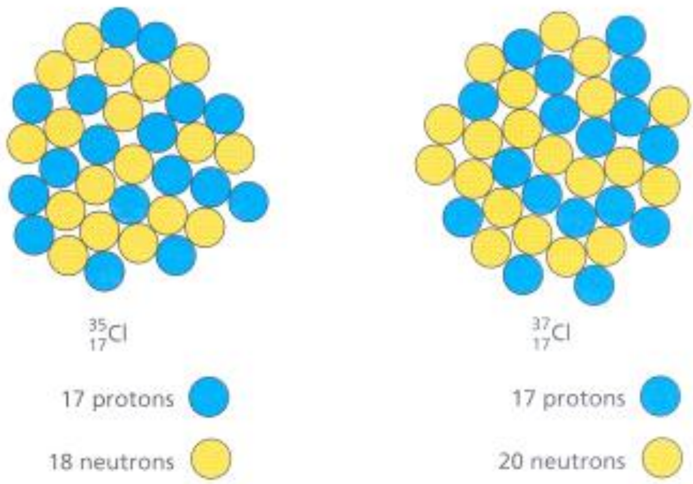


Figure 3.4 The two isotopes of chlorine.

Some of the atoms of certain isotopes are unstable because of the extra number of neutrons, and they are said to be radioactive. The best known elements which have radioactive isotopes are uranium and carbon. The major isotopes of these elements are shown in Table 3.4.

Table 3.4 Some atoms and their isotopes.

Element	Symbol	Particles present
Hydrogen (Deuterium) (Tritium)	$^{1}_{1}H$	1 e, 1 p, 0 n
	$^{2}_{1}H$	1 e, 1 p, 1 n
	$^{3}_{1}H$	1 e, 1 p, 2 n
Carbon	$^{12}_{6}C$	6 e, 6 p, 6 n
	$^{13}_{6}C$	6 e, 6 p, 7 n
	$^{14}_{6}C$	6 e, 6 p, 8 n
Oxygen	$^{16}_{8}O$	8 e, 8 p, 8 n
	$^{17}_{8}O$	8 e, 8 p, 9 n
	$^{18}_{6}O$	8 e, 8 p, 10 n
Strontium	$^{86}_{38}Sr$	38 e, 38 p, 48 n
	$^{88}_{38}Sr$	38 e, 38 p, 50 n
	$^{90}_{38}Sr$	38 e, 38 p, 52 n
Uranium	$^{235}_{92}U$	92 e, 92 p, 143 n

	$^{238}_{92}U$	92 e, 92 p, 146 n

The mass spectrometer

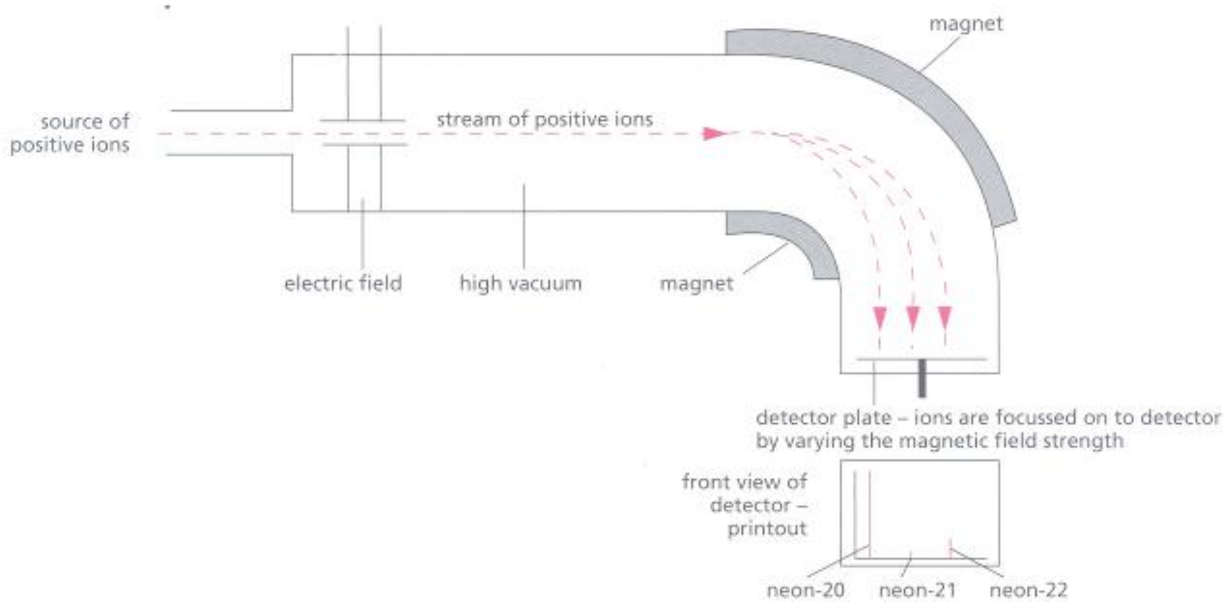


Figure 3.5 A diagram of a mass spectrometer.

How do we know isotopes exist? They were first discovered by scientists using apparatus called a mass spectrometer (Figure 3.5). The first mass spectrometer was built by the British scientist Francis Aston in 1919 and enabled scientists to compare the relative masses of atoms accurately for the first time.

A vacuum exists inside a mass spectrometer. A sample of the vapour of the element is injected into a chamber where it is bombarded by electrons. The subsequent collisions that take place cause the atoms in the vapour to lose one of their electrons and so form positive ions. The beam of positive ions is accelerated by an electric field and then deflected by a magnetic field. The amount of deflection depends on the different masses of the positive ions. The lighter ions, which were formed from the lighter isotopes, are deflected more than the heavier ones. In this way particles with different masses can be separated and identified. A detector counts the number of each of the ions that fall upon it and so a measure of the percentage abundance of each isotope is obtained. A typical mass spectrum for chlorine is shown in Figure 3.6.

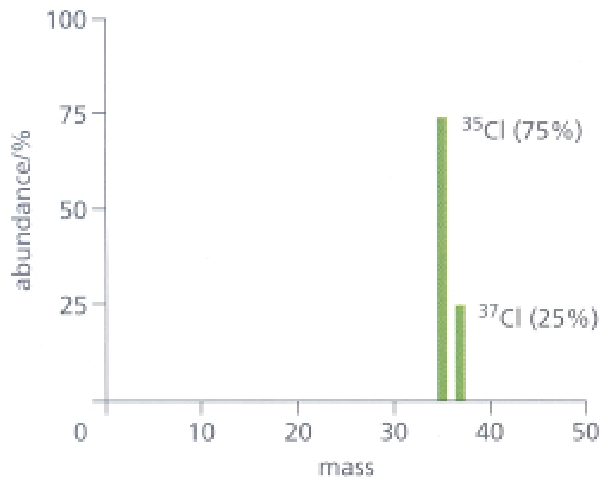


Figure 3.6 The mass spectrum for chlorine.

Relative atomic mass

The average mass of a large number of atoms of an element is called its **relative atomic mass** (symbol $\mathbf{A_r}$). This quantity takes into account the percentage abundance of all the isotopes of an element which exist.

In 1961 the International Union of Pure and Applied Chemistry (IUPAC) recommended that the standard used for the A_i scale was carbon-12. An atom of carbon-12 was taken to have a mass of 12 amu. The A,. of an element is now defined as the average mass of its isotopes compared with one-twelfth the mass of one atom of carbon-12:

$$A_r = \frac{\text{average mass of isotopes of the element}}{\frac{1}{12} \times \text{mass of 1 atom of carbon-12}}$$

Note: $\frac{1}{12}$ of the mass of one carbon-12 atom = 1 amu.

For example, chlorine has two isotopes:

	$^{35}_{17}Cl$	$^{37}_{17}Cl$
% Abundance	75	25

Hence the 'average mass' of a chlorine atom is:

$$\frac{(75\times 35)+(25\times 37)}{100}=35.5$$

$A_{\mathbf{r}}$ = 35.5 amu

Questions

1. Calculate the number of neutrons in the following atoms:

a. $^{27}_{13}Al$ b. $^{31}_{15}P$ c. $^{262}_{107}Uns$ d. $^{190}_{76}Os$

2. Given that the percentage abundance of $^{20}_{10}Ne$ is 90% and that of $^{22}_{10}Ne$ is 10%, calculate the A_r of neon.

The arrangement of electrons in atoms

The nucleus of an atom contains the heavier sub-atomic particles — the protons and the neutrons. The electrons, the lightest of the sub-atomic particles, move around the nucleus at great distances from the nucleus relative to their size. They move very fast in electron energy levels very much as the planets orbit the Sun.

It is not possible to give the exact position of an electron in an energy level. However, we can state that electrons can only occupy certain, definite energy levels and that they cannot exist between them. Each of the electron energy levels can hold only a certain number of electrons.

- First energy level holds up to 2 electrons.
- Second energy level holds up to 8 electrons.
- Third energy level holds up to 18 electrons.

There are further energy levels which contain increasing numbers of electrons.

The third energy level can be occupied by a maxi-mum of 18 electrons. However, when eight electrons have occupied this level a certain stability is given to the atom and the next two electrons go into the fourth energy level, and then the remaining ten electrons complete the third energy level.

The electrons fill the energy levels starting from the energy level nearest to the nucleus, which has the lowest energy. When this is full (with two electrons) the next electron goes into the second energy level. When this energy level is full with eight electrons, then the electrons begin to fill the third and fourth energy levels as stated above.

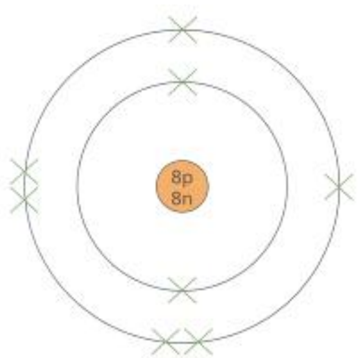


Figure 3.7 Arrangement of electrons in an oxygen atom.

For example, a $^{16}_{8}O$ atom has an atomic number of 8 and therefore has eight electrons. Two of the eight electrons enter the first energy level, leaving six to occupy the second energy level, as shown in Figure 3.7. The electron configuration for oxygen can be written in a shorthand way as 2,6.

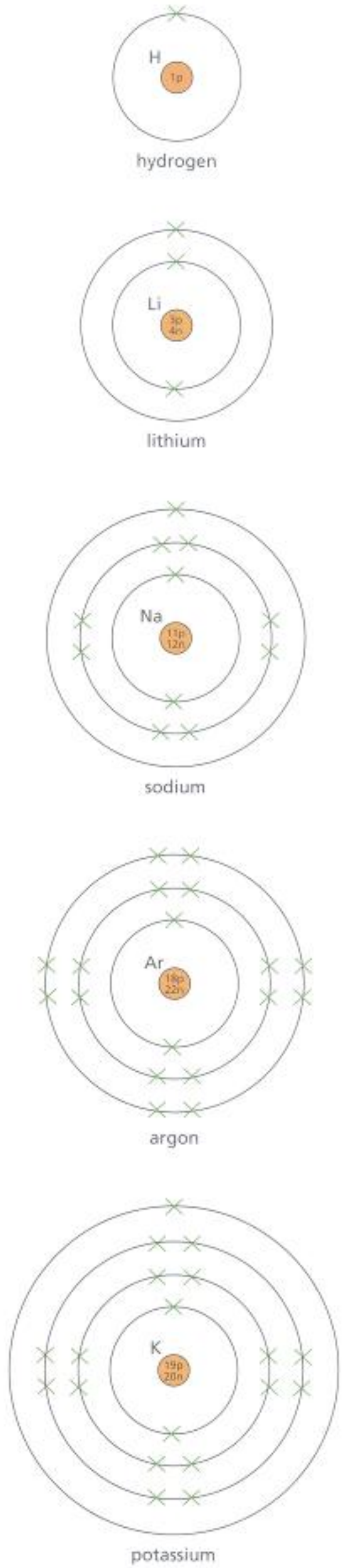


Figure 3.8 Electron arrangements of hydrogen, lithium, sodium, argon and potassium.

There are 115 elements, and Table 3.5 shows the way in which the electrons are arranged in the first 20 of these elements. The way in which the electrons are distributed is called the **electron structure** or **electron configuration**. Figure 3.8 shows the electron configuration of a selection of atoms.

Table 3.5 Electron arrangement in the first 20 elements.

Element	Symbol	Atomic Number	Number of electrons	Electron structure
Hydrogen	H	1	1	1
Helium	He	2	2	2
Lithium	Li	3	3	2,1
Beryllium	Be	4	4	2,2
Boron	B	5	5	2,3
Carbon	C	6	6	2,4
Nitrogen	N	7	7	2,5
Oxygen	O	8	8	2,6
Fluorine	F	9	9	2,7
Neon	Ne	10	10	2,8
Sodium	Na	11	11	2,8,1
Magnesium	Mg	12	12	2,8,2
Aluminium	Al	13	13	2,8,3
Silicon	Si	14	14	2,8,4
Phosphorus	P	15	15	2,8,5
Sulphur	S	16	16	2,8,6
Chlorine	Cl	17	17	2,8,7
Argon	Ar	18	18	2,8,8
Potassium	K	19	19	2,8,8,1
Calcium	Ca	20	20	2,8,8,2

Questions

1. How many electrons may be accommodated in the first three energy levels?

2. What is the same about the electron structures of:

a. lithium, sodium and potassium?
b. beryllium, magnesium and calcium?

The periodic table

The periodic table was devised in 1869 by the Russian Dimitri Mendeleev, who was the Professor of Chemistry at St Petersburg University (Figure 3.9). His periodic table was based on the chemical and physical properties of the 63 elements that had been discovered at that time.

Figure 3.9 Dmitri Mendeleev (1834–1907).

However, other scientists had also attempted to categorise the known elements. In 1817, Johann Dobereiner noticed that the atomic weight (now called atomic mass) of strontium fell midway between the weights of calcium and barium. These were elements which possessed similar chemical properties. They formed a **triad** of elements. Other triads were also discovered, composed of:

chlorine, bromine, iodine

lithium, sodium, potassium

He called this the **'Law of Triads'**. This encouraged other scientists to

search for patterns.

In 1865, John Newlands, an English chemist, arranged the 56 known elements in order of increasing atomic weight. He realised when he did this that every eighth element in the series was similar.

H **Li** Be B C N O F **Na** Mg Al Si P S Cl **K**

He likened this to music and called it the **'Law of Octaves'**. It fell down, however, because some of the weights were inaccurate and there were elements that had not been discovered then.

Period	**Group**							
	1	2	3	4	5	6	7	8
1	H							
2	Li	Be	B	C	N	O	F	
3	Na	Mg	Al	Si	P	S	Cl	
4	K Cu	Ca Zn	* *	Ti †	V As	Cr Se	Mn Br	Fe Co Ni

Figure 3.10 Mendeleev's periodic table. He left gaps for undiscovered elements.

Mendeleev's classification proved to be the most successful. Mendeleev arranged all the 63 known elements in order of increasing atomic weight but in such a way that elements with similar properties were in the same vertical column. He called the vertical columns **groups** and the horizontal rows **periods** (Figure 3.10). If necessary he left gaps in the table.

As a scientific idea, Mendeleev's periodic table was tested by making predictions about elements that were unknown at that time but could possibly fill the gaps.

Three of these gaps are shown by the symbols * and † in Figure 3.10. As new elements were discovered, they were found to fit easily into the classification. For example, Mendeleev predicted the properties of the missing element 'eka-silicon' (†). He predicted the colour, density and melting point as well as its atomic weight.

In 1886 the element we now know as germanium was discovered in Germany by Clemens Winkler; its properties were almost exactly those

Mendeleev had predicted. In all, Mendeleev predicted the atomic weight of ten new elements, of which seven were eventually discovered — the other three, atomic weights 45, 146 and 175 do not exist!

The success of Mendeleev's predictions showed that his ideas were probably correct. His periodic table was quickly accepted by scientists as an important summary of the properties of the elements.

Mendeleev's periodic table has been modified in the light of work carried out by Rutherford and Moseley. Discoveries about sub-atomic particles led them to realise that the elements should be arranged by atomic number. In the modern periodic table the 115 known elements are arranged in order of increasing atomic number (Figure 3.11).

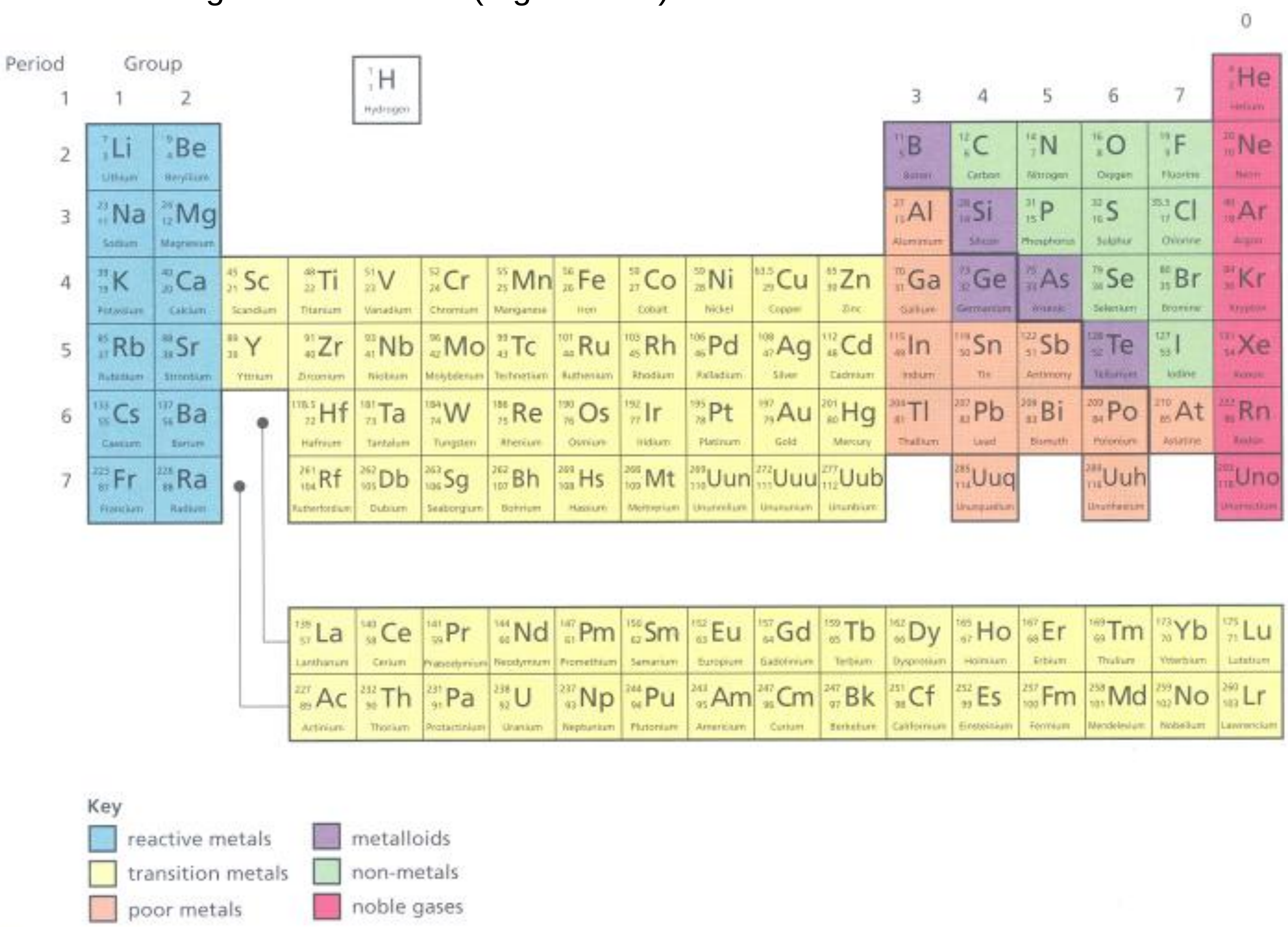


Figure 3.11 The modern periodic table.

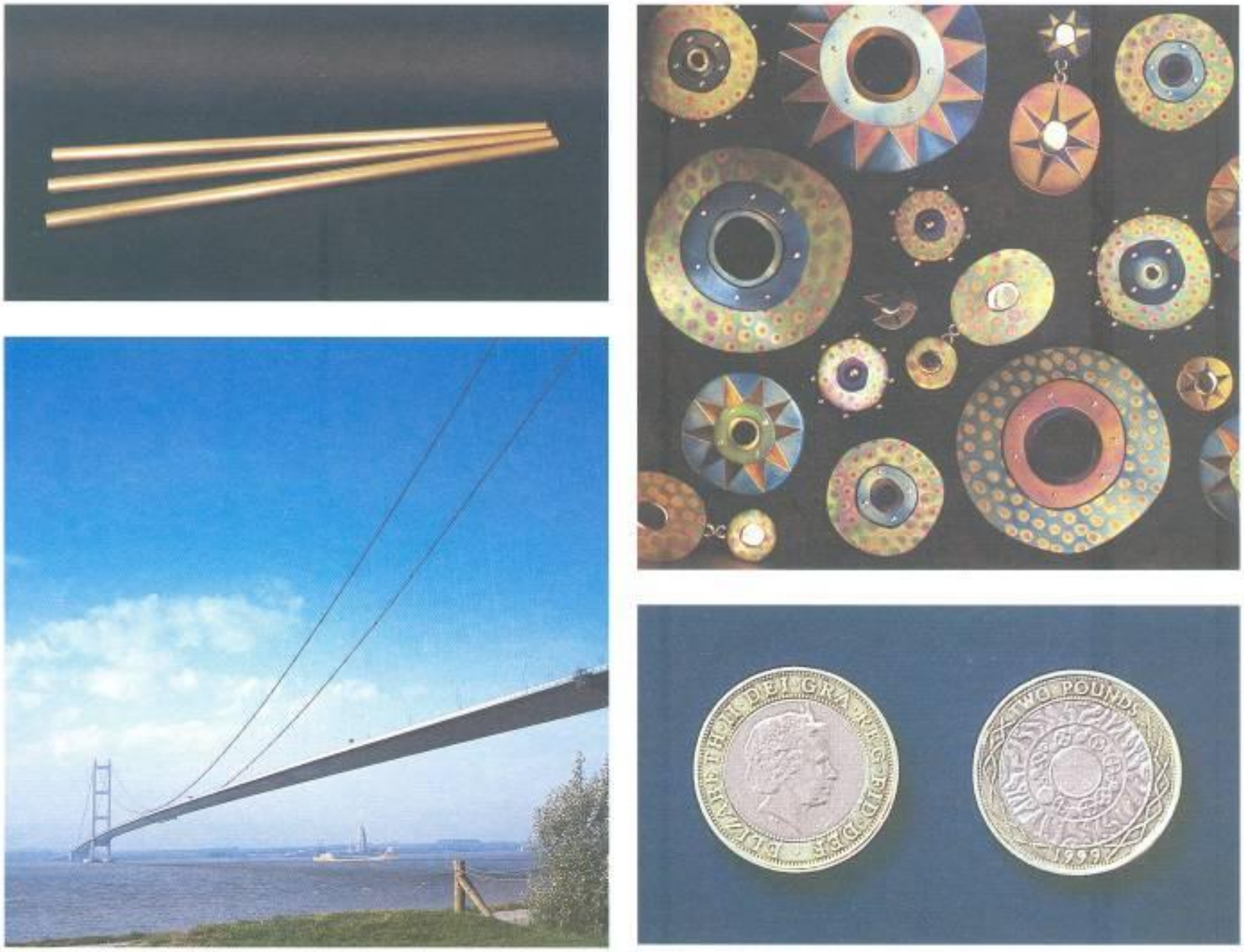


Figure 3.12 Transition elements have a wide range of uses, both as elements and as alloys.

Those elements with similar chemical properties are found in the same columns or **groups**. There are eight groups of elements. The first column is called group 1; the second group 2; and so on up to group 7. The final column in the periodic table is called group 0. Some of the groups have been given names.

Group 1: The alkali metals
Group 2: The alkaline earth metals
Group 7: The halogens
Group 0: Inert gases or noble gases

The horizontal rows are called **periods** and these are numbered 1-7 going down the periodic table.

Between groups 2 and 3 is the block of elements known as the transition elements (Figure 3.12).

The periodic table can be divided into two as shown by the bold line that starts beneath boron, opposite. The elements to the left of this line are metals (fewer than three-quarters) and those on the right are non-metals (fewer then one-quarter). The elements which lie on this dividing line are known as metalloids (Figure 3.13). These elements behave in some ways as metals and in others as non-metals.

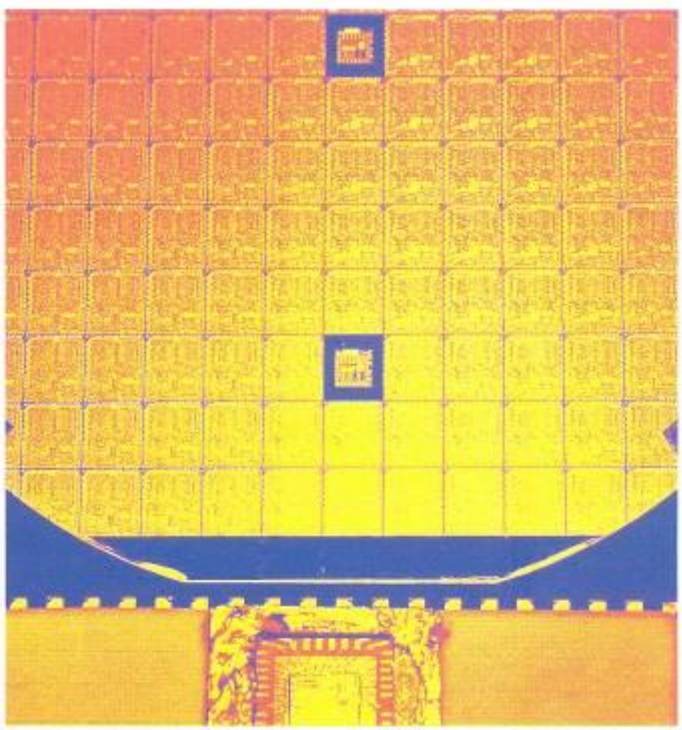


Figure 3.13 The metalloid silicon is used to make silicon 'chips'.

Electron configuration and the periodic table

Now that the number of electrons in the outer energy level has been established, it can be seen that it corresponds with the number of the group in the periodic table in which the element is found. For example, the elements shown in Table 3.6 have one electron in their outer energy level and they are all found in group 1. The elements in group 0, however, are an exception to this rule, as they have two or eight electrons in their outer energy level. The outer electrons are mainly responsible for the chemical

properties of any element, and, therefore, elements in the same group have similar chemical properties (Tables 3.7 and 3.8).

Table 3.6 Electron configuration of the first three elements of group 1.

Element	**Symbol**	**Atomic number**	**Electron configuration**
Lithium	Li	3	2,1
Sodium	Na	11	2,8,1
Potassium	K	19	2,8,8,1

Table 3.7 Electron configuration of the first three elements of group 2.

Element	**Symbol**	**Atomic number**	**Electron configuration**
Beryllium	Be	4	2,2
Magnesium	Mg	12	2,8,2
Calcium	Ca	20	2,8,8,2

Table 3.8 Electron configuration of the first three elements in group 7.

Element	**Symbol**	**Atomic number**	**Electron configuration**
Fluorine	F	9	2,7
Chlorine	Cl	17	2,8,7
Bromine	Br	35	2,8,18,7

Group 1 - the alkali metals

Group 1 consists of the five metals lithium, sodium, potassium, rubidium and caesium, and the radioactive element francium. Lithium, sodium and potassium are commonly available for use in school. They are all very reactive metals and they are stored under oil to prevent them coming into contact with water or air. These three metals have the following properties.

- They are good conductors of electricity and heat.
- They are soft metals.
- They are metals with low densities.
- They have shiny surfaces when freshly cut with a knife (Figure 3.14).

Figure 3.14 Freshly cut sodium.

- They burn in oxygen or air, with characteristic flame colours, to form white solid oxides. For example, lithium reacts with the oxygen in the air to form white lithium oxide, according to the following equation:

$$\text{lithium} + \text{oxygen} \rightarrow \text{lithium oxide}$$

$$4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$$

These group 1 oxides all dissolve in water to form alkaline solutions of the metal hydroxide.

$$\text{lithium oxide} + \text{water} \rightarrow \text{lithium hydroxide}$$

$$Li_2O(s) + H_2O(l) \rightarrow 2LiOH(aq)$$

- They react vigorously with water to give an alkaline solution of the metal hydroxide as well as producing hydrogen gas.

For example:

$$\text{potassium} + \text{water} \rightarrow \text{potassium hydroxide} + \text{hydrogen gas}$$

$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$$

Of these three metals, potassium is the most reactive towards water (Figure 3.15), followed by sodium and then lithium. Such gradual changes we call **trends**. Trends are useful to chemists as they allow predictions to be made about elements we have not observed in action.

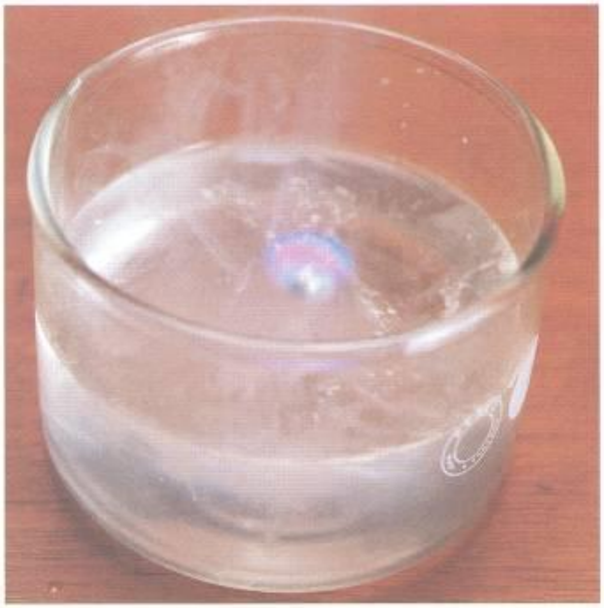
- They react vigorously with halogens, such as chlorine, to form metal halides, for example sodium chloride (Figure 3.16).

$$\text{sodium} + \text{chorine} \rightarrow \text{sodium chloride}$$

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

Considering the group as a whole, the further down the group you go the more reactive the metals become. Francium is, therefore, the most reactive of the group 1 metals.

Table 3.6 shows the electron configuration of the first three elements of group 1. You will notice in each case that the outer energy level contains only one electron. When these elements react they lose this outer electron, and in doing so become more stable, because they obtain the electron configuration of a noble gas. You will learn more about the stable nature of these gases later in this chapter.



a Potassium reacts very vigorously with cold water.

b An alkaline solution is produced when potassium reacts with water.

Figure 3.15

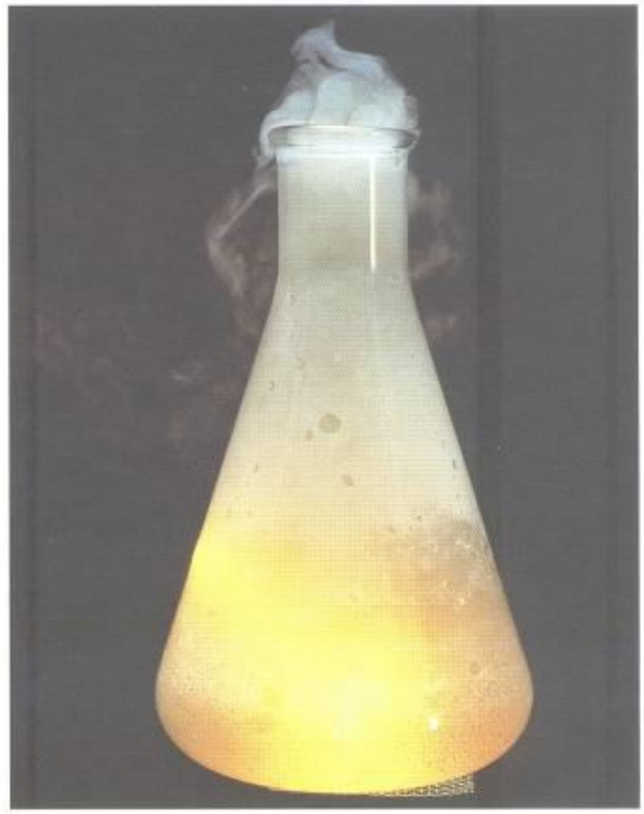


Figure 3.16 A very vigorous reaction takes place when sodium burns in chlorine gas. Sodium chloride is produced.

When, for example, the element sodium reacts it loses its outer electron. This requires energy to over-come the electrostatic attractive forces between the outer electron and the positive nucleus (Figure 3.17).

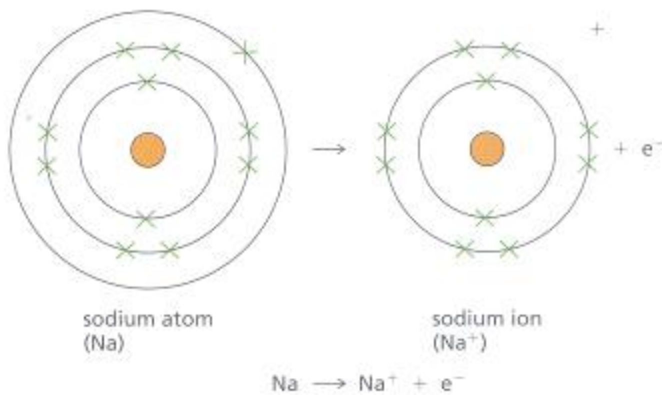


Figure 3.17 This sodium atom loses an electron to become a sodium ion.

Look at Figure 3.18. Why do you think potassium is more reactive than lithium or sodium?

Potassium is more reactive because less energy is required to remove the outer electron from its atom than for lithium or sodium. This is because as you go down the group the size of the atoms increases and the outer electron gets further away from the nucleus and becomes easier to remove.

Questions

1. Write word and balanced chemical equations for the reactions between:

a. sodium and oxygen
b. sodium and water.

2. Write word and balanced chemical equations for the reactions between:

a. magnesium and water
b. calcium and oxygen.

3. Account for the fact that calcium is more reactive than magnesium.

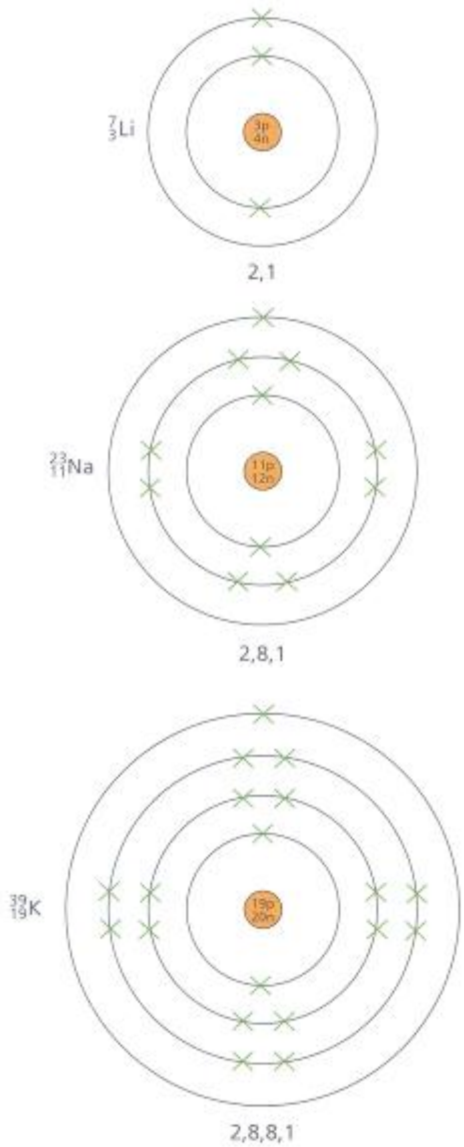


Figure 3.18 Electron structures of lithium, sodium and potassium.

Group 2 - the alkaline earth metals

Group 2 consists of the five metals beryllium, magnesium, calcium, strontium and barium, and the radio-active element radium. Magnesium and calcium are generally available for use in school. These metals have the following properties.

- They are harder than those in group 1.

- They are silvery-grey in colour when pure and clean. They tarnish

quickly, however, when left in air due to the formation of a metal oxide on their surfaces (Figure 3.19).

- They are good conductors of heat and electricity.

- They burn in oxygen or air with characteristic flame colours to form solid white oxides. For example:

$$\text{magnesium} + \text{oxygen} \rightarrow \text{magnesium oxide}$$

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

- They react with water, but they do so much less vigorously than the elements in group 1. For example:

$$\text{calcium} + \text{water} \rightarrow \text{calcium hydroxide} + \text{hydrogen gas}$$

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

Considering the group as a whole, the further down the group you go, the more reactive the elements become.

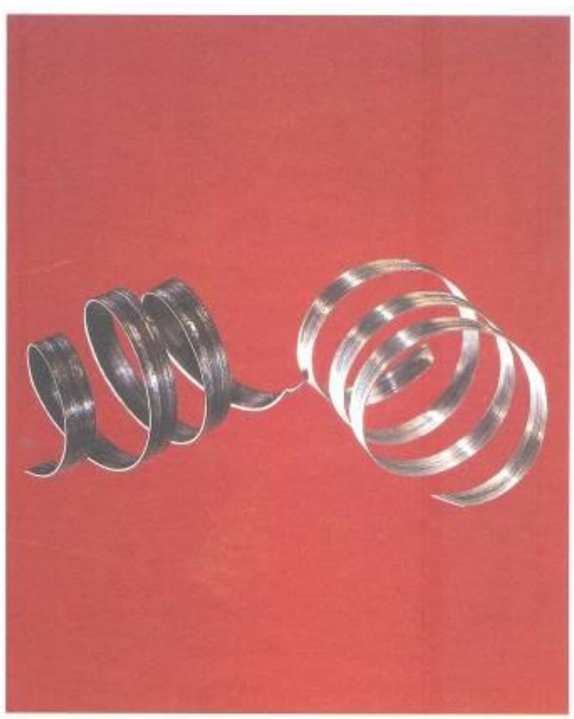


Figure 3.19 Tarnished (left) and cleaned-up magnesium.

Flame colours

If a clean nichrome wire is dipped into a metal compound and then held in the hot part of a Bunsen flame, the flame can become coloured (Figure 3.20). Certain metal ions may be detected in their compounds by observing their flame colours (Table 3.9).

Table 3.9 Characteristic flame colours of some metal ions.

	Metal	**Flame colour**
Group 1	Lithium	Crimson
	Sodium	Golden yellow
	Potassium	Lilac
	Rubidium	Red
	Caesium	Blue
Group 2	Calcium	Brick red
	Strontium	Crimson
	Barium	Apple green
Others	Lead	Blue-white
	Copper	Green

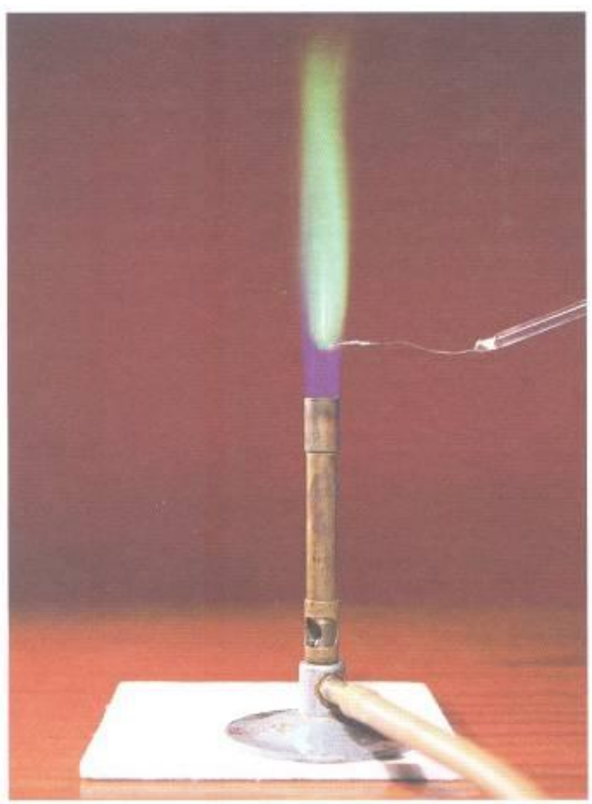


Figure 3.20 The green colour is characteristic of copper.

A flame colour is obtained as a result of the electrons in the particular ions being excited when they absorb energy from the flame which is then emitted as visible light. The different electron configurations of the different ions, therefore, give rise to the different colours (Figure 3.21).

Figure 3.21 Different metal ions are used to produce the colours seen in fireworks.

Group 7 - the halogens

Group 7 consists of the four elements fluorine, chlorine, bromine and iodine, and the radioactive element astatine. Of these five elements, chlorine, bromine and iodine are generally available for use in school.

- These elements are coloured and darken going down the group (Table 3.10).

- They exist as diatomic molecules, for example Cl_2, Br_2 and I_2.

- They show a gradual change from a gas (Cl_2), through a liquid (Br_2), to a solid (I_2) (Figure 3.22).

- They form molecular compounds with other non-metallic elements, for example HCl.

- They react with hydrogen to produce the hydrogen halides, which

dissolve in water to form acidic solutions (pH < 7).

hydrogen + chlorine → hydrogen chloride

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

hydrogen chloride + water → hydrochloric acid

$$HCl(g) + H_2O \rightarrow HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

- They react with metals to produce ionic metal halides, for example chlorine and iron produce iron chloride.

iron + chlorine → iron(III) chloride

$$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$$

To find out about the extraction of bromine, see Chapter 10.

Table 3.10 Colours of some halogens.

Halogen	**Colour**
Chlorine	Pale green
Bromine	Red-brown
Iodine	Purple-black

a Chlorine, bromine and iodine.

b Chlorine gas bleaches moist indicator paper.

Figure 3.22

Displacement reactions

If chlorine is bubbled into a solution of potassium bromide the less reactive halogen, bromine, is displaced by the more reactive halogen, chlorine, as you can see from Figure 3.23:

potassium bromide + chlorine $\rightarrow$ potassium chloride + bromine

$$2KBr(aq) + Cl_2(g) \rightarrow 2KCl(aq) + Br_2(aq)$$

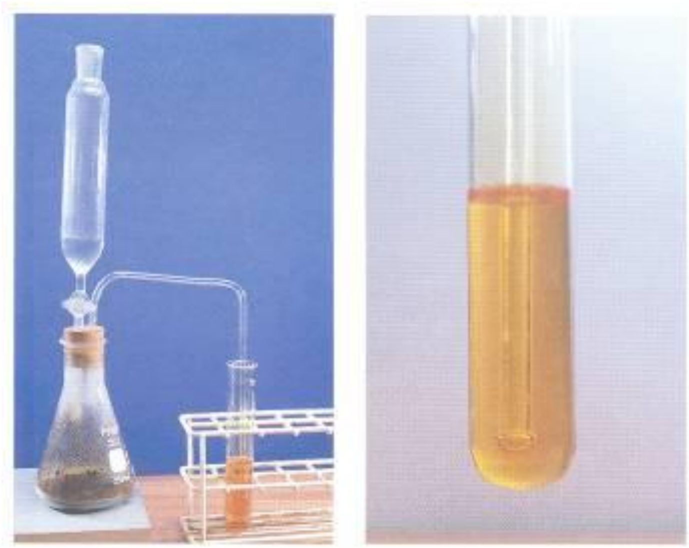


Figure 3.23 Bromine being displaced.

The observed order of reactivity of the halogens, con-firmed by similar displacement reactions, is:

$$\xrightarrow{\text{Decreasing reactivity}}$$

chlorine bromine iodine

You will notice that, in contrast to the elements of groups 1 and 2, the order of reactivity decreases on going down the group.

Table 3.11 shows the electron configuration for chlorine and bromine. In each case the outer energy level contains seven electrons. When these elements react they gain one electron per atom to gain the stable electron configuration of a noble gas. You will learn more about the stable nature of these gases in the next section. For example, when chlorine reacts it gains

a single electron and forms a negative ion (Figure 3.24).

Table 3.11 Electron configuration of chlorine and bromine.

Element	Symbol	Atomic number	Electron configuration
Chlorine	Cl	17	2,8,7
Bromine	Br	35	2,8,18,7

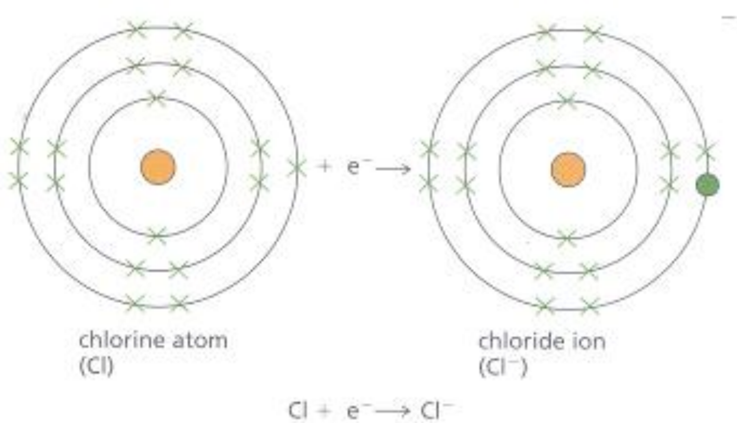


Figure 3.24 A chlorine atom gains an electron to form a chloride ion.

Chlorine is more reactive than bromine because the incoming electron is being more strongly attracted into the outer energy level of the smaller atom. The attractive force on it will be greater than in the case of bromine, since the outer energy level of chlorine is closer to the nucleus. As you go down the group this outermost extra electron is further from the nucleus. It will, therefore, be held less securely, and the resulting reactivity of the elements in group 7 will decrease down the group.

The halogens and halogenic compounds are used in a multitude of different ways (Figure 3.25).

Figure 3.25 The halogens have many varied uses.

- Fluorine is used in the form of fluorides in drinking water and toothpaste because it reduces tooth decay by hardening the enamel on teeth.

- Chlorine is used to make PVC plastic as well as household bleaches. It is also used to kill bacteria and viruses in drinking water (Chapter 10).

- Bromine (Chapter 10) is used to make disinfectants, medicines and fire retardants.

- Iodine is used in medicines and disinfectants and also as a photographic chemical.

Questions

1. Write word and balanced chemical equations for the reactions between:

a. bromine and potassium iodide solution
b. bromine and potassium chloride solution.

If no reaction will take place, write 'no reaction' and explain why.

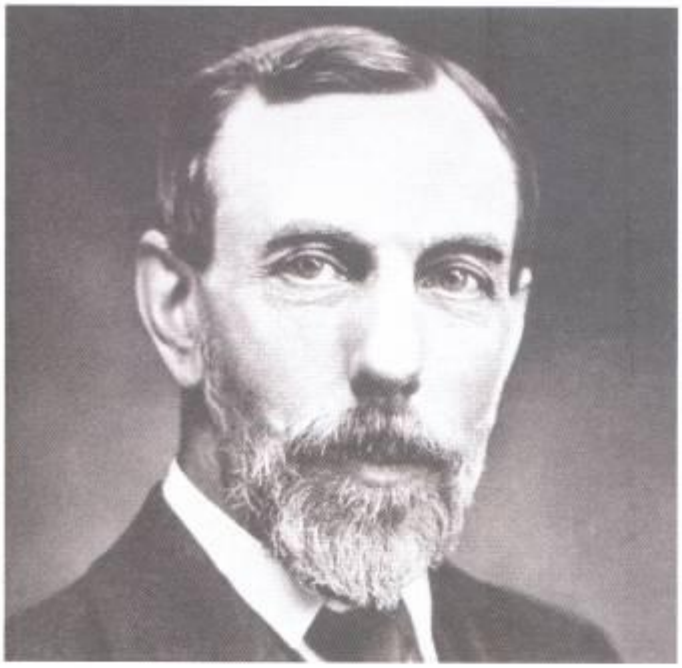
2. Write down the names and symbols for the noble gases not given in Table 3.12 and use your research skills to find a use for each.

Group 0 - the noble gases

Helium, neon, argon, krypton, xenon and the radio-active element radon make up a most unusual group of non-metals, called the noble gases. They were all discovered after Mendeleev had published his periodic table. They were discovered between 1894 and 1900, mainly through the work of the British scientists Sir William Ramsay (Figure 3.26a) and Lord John William Strutt Rayleigh (Figure 3.26b).

- They are colourless gases.

- They exist as individual atoms, for example He, Ne and Ar.
- They are very unreactive.



a Sir William Ramsay (1852–1916).

b Lord Rayleigh (1842–1919).

Figure 3.26 Both helped to discover the noble gases and won the Nobel Prize in Chemistry in 1904 for their work.

No compounds of helium, neon or argon have ever been found. However, more recently a number of compounds of xenon and krypton with fluorine and oxygen have been produced, for example XeF_6.

These gases are chemically unreactive because they have electron configurations which are stable and very difficult to change (Table 3.12). They are so stable that other elements attempt to attain these electron configurations during chemical reactions (Chapter 4). You have probably seen this in your study of the elements of groups 1, 2 and 7.

Although unreactive, they have many uses. Argon, for example, is the gas used to fill light bulbs to prevent the tungsten filament reacting with air. Neon is used extensively in advertising signs and in lasers. Further uses of these gases are discussed in Chapter 10.

Table 3.12 Electron configuration of helium, neon and argon.

Element	Symbol	Atomic number	Electron configuration
Helium	He	2	2
Neon	Ne	10	2,8
Argon	Ar	18	2,8,8

Helium is separated from natural gas by the liquefaction of the other gases. The other noble gases are obtained in large quantities by the fractional distillation of liquid air (Chapter 10).

Transition elements

This block of metals includes many you will be familiar with, for example copper, iron, nickel, zinc and chromium (Figure 3.27).

- They are harder and stronger than the metals in groups 1 and 2.
- They have much higher densities than the metals in groups 1 and 2.
- They have high melting points (except for mercury, which is a liquid at room temperature).
- They are less reactive metals.
- They form a range of brightly coloured compounds (Figure 3.28).
- They are good conductors of heat and electricity.
- They show catalytic activity (Chapter 11) as elements and compounds. For example, iron is used in the industrial production of ammonia gas (Haber process, Chapter 15).
- They do not react (corrode) so quickly with oxygen and/or water.
- They form more than one simple ion. For example, copper forms Cu^{+} and Cu^{2+} in compounds such as Cu_2O and $CuSO_4$; iron forms Fe^{2+} and Fe^{3+} in compounds such as $FeSO_4$ and $FeCl_3$.

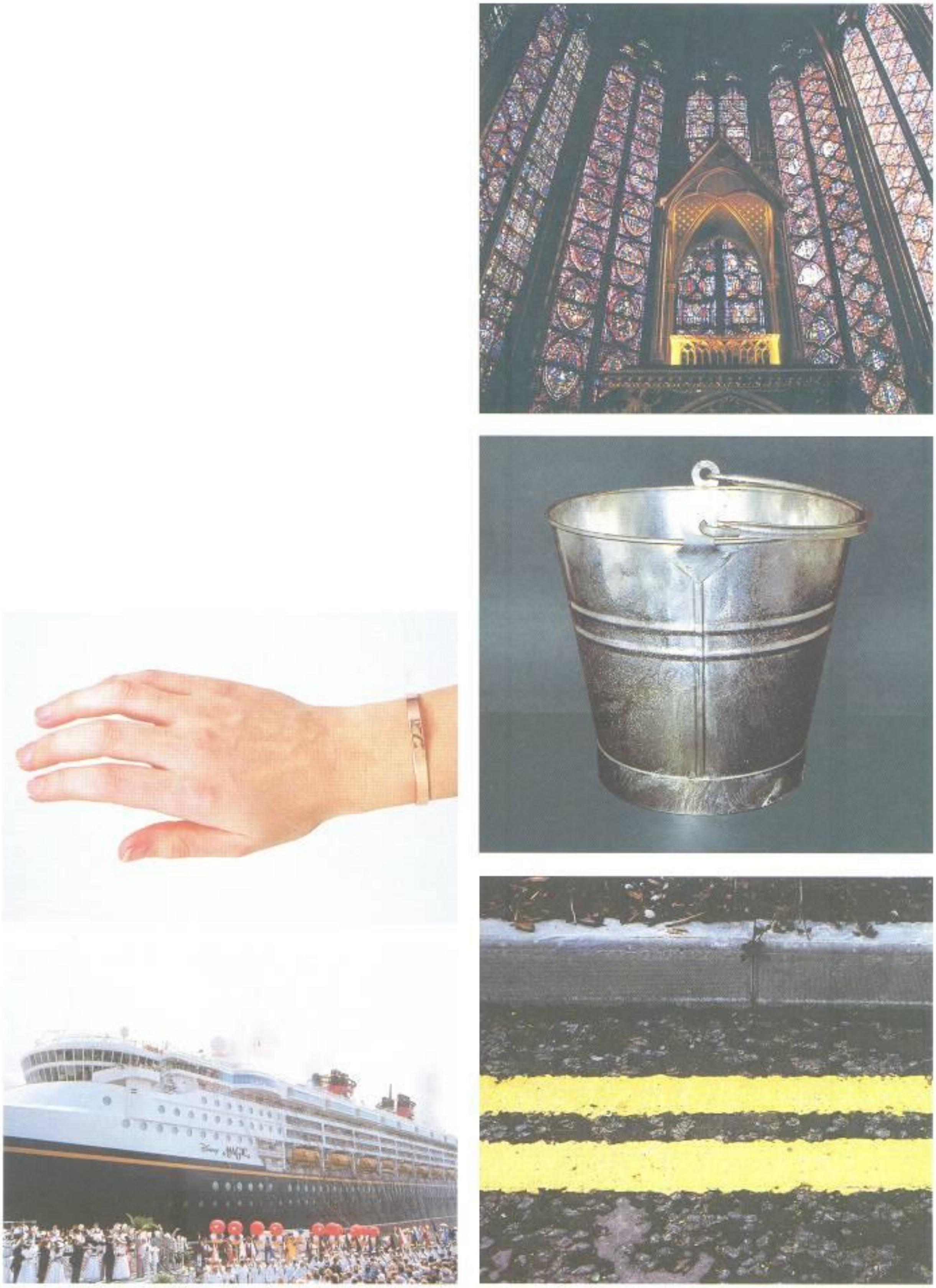


Figure 3.27 Everyday uses of transition elements and their compounds. They are often known as the 'everyday metals'.

a Some solutions of coloured transition element compounds.

b The coloured compounds of transition elements can be seen in these pottery glazes.

Figure 3.28

Questions

1. Look at the photographs in Figure 3.27 and decide which properties are important when considering the particular use the metal is being put to.

2. Which groups in the periodic table contain:

a. only metals?
b. only non-metals?
c. both metals and non-metals?

The position of hydrogen

Hydrogen is often placed by itself in the periodic table. This is because the properties of hydrogen are unique. However, profitable comparisons can be made with the other elements. It is often shown at the top of either group 1 or group 7, but it cannot fit easily into the trends shown by either group, see Table 3.13.

Table 3.13 Comparison of hydrogen with lithium and fluorine.

Lithium	**Hydrogen**	**Fluorine**
Solid	Gas	Gas
Forms a positive ion	Forms positive or negative ions	Forms a negative ion
1 electron in outer energy level	1 electron in outer energy level	1 electron short of a full outer energy level
Loses 1 electron to form a noble gas configuration	Needs 1 electron to form a noble gas configuration	Needs 1 electron to form a noble gas configuration

Checklist

After studying Chapter 3 you should know and understand the following terms.

Atomic mass unit Exactly $\frac{1}{12}$ of the mass of one atom of the most abundant isotope of carbon-12.

Atomic number (proton number) Symbol Z. The number of protons in the nucleus of an atom. The number of electrons present in an atom. The order of the element within the periodic table.

Displacement reaction A reaction in which a more reactive element displaces a less reactive element from solution.

Electron A fundamental sub-atomic particle with a negative charge present in all atoms within energy levels around the nucleus.

Electron configuration A shorthand method of describing the arrangement of electrons within the energy levels of an atom.

Electron energy levels The allowed energies of electrons in atoms.

Electrostatic force of attraction A strong force of attraction between opposite charges.

Group A vertical column of the periodic table containing elements with similar properties with the same number of electrons in their outer energy levels. They have an increasing number of inner energy levels as you descend the group.

Ion An atom or group of atoms which has either lost one or more electrons, making it positively charged, or gained one or more electrons, making it negatively charged.

Isotopes Atoms of the same element which possess different numbers of neutrons. They differ in mass number (nucleon number).

Mass number (nucleon number) Symbol A. The total number of protons

and neutrons found in the nucleus of an atom.

Mass spectrometer A device in which atoms or molecules are ionised and then accelerated. Ions are separated according to their mass.

Metalloid (semi-metal) Any of the class of chemical elements intermediate in properties between metals and non-metals, for example boron and silicon.

Neutron A fundamental, uncharged sub-atomic particle present in the nuclei of atoms.

Periodic table A table of elements arranged in order of increasing atomic number to show the similarities of the chemical elements with related electron configurations.

Periods Horizontal rows of the periodic table. Within a period the atoms of all the elements have the same number of occupied energy levels but have an increasing number of electrons in the outer energy level.

Proton A fundamental sub-atomic particle which has a positive charge equal in magnitude to that of an electron. Protons occur in all nuclei.

Relative atomic mass Symbol A_r

$$A_r = \frac{\text{average mass of isotopes of the element}}{\frac{1}{12} \times \text{mass of 1 atom of carbon-12}}$$

Atomic structure and the periodic table

Additional questions

1. An atom X has an atomic number of 19 and relative atomic mass of 39.

a. How many electrons, protons and neutrons are there in an atom of X?

b. How many electrons will there be in the outer energy level (shell) of an atom of X?

c. Write down the symbol for the ion X will form.

d. Which group of the periodic table would X be in?

e. (i) How would you expect X to react with water?

(ii) Write a word and balanced chemical equation for this reaction.

2. The atomic number of barium (Ba) is 56. It is in group 2 of the periodic table.

a. How many electrons would you expect a barium atom to contain in its outer energy level?

b. How would you expect barium to react with chlorine? Write a word and balanced chemical equation for this reaction.

c. How would you expect barium to react with water? Write a word and balanced chemical equation for this reaction.

d. Write down the formulae of the bromide and sulphate of barium.

3. Find the element germanium (Ge) in the periodic table.

a. Which group of the periodic table is this element in?

b. How many electrons will it have in its outer energy level (shell)?

c. Is germanium a metal or a non-metal?

d. What is the formula of the chloride of germanium?

e. Name and give the symbols of the other elements in this group.

4. Three members of the halogens are: $^{35.5}_{17}Cl$, $^{80}_{35}Br$ and $^{127}_{53}I$.

a (i) Write down the electron structure of an atom of chlorine.

(ii) Why is the relative atomic mass of chlorine not a whole number?

(iii) How many protons are there in an atom of bromine?

(iv) How many neutrons are there in an atom of iodine?

(v) State and account for the order of reactivity of these elements.

b. When potassium is allowed to burn in a gas jar of chlorine, in a fume cupboard, clouds of white smoke are produced.

(i) Why is this reaction carried out in a fume cupboard?

(ii) What does the white smoke consist of?

(iii) Write a word and balanced chemical equation for this reaction.

(iv) Describe what you would expect to see when potassium is allowed to burn safely in a gas jar of bromine vapour. Write a word and balanced chemical equation for this reaction.

5. 'By using displacement reactions it is possible to deduce the order of reactivity of the halogens.' Discuss this statement with reference to the elements bromine, iodine and chlorine only.

6. Use the information given in the table below to answer the questions below concerning the elements Q, R, S, T and X.

Element	Atomic number	Mass number	Electron structure
Q	3	7	2,1
R	20	40	2,8,8,2
S	18	40	2,8,8
T	8	18	2,6
X	19	39	2,8,8,1

a. Which element has 22 neutrons in each atom?

b. Which element is a noble gas?

c. Which two elements form ions with the same electron structure as neon?

d. Which two elements are in the same group of the periodic table and which group is this?

e. Place the elements in the table above into the periods in which they belong.

f. Which is the most reactive metal element shown in the table?

g. (i) Which of the above elements is calcium?

(ii) What colour flame would compounds containing calcium produce?

7a. $^{69}_{31}Ga$ and $^{71}_{31}Ga$ are isotopes of gallium.

With reference to this example, explain what you understand by the term isotope.

b. A sample of gallium contains 60% of atoms of $^{69}_{31}Ga$ and 40% of atoms of $^{71}_{31}Ga$. Calculate the relative atomic mass of this sample of gallium.

8. Copy and complete the following table with reference to the periodic table.

Element name	**Symbol**	**Atomic number**	**Mass number**	**Number of neutrons**	$^{A}_{Z}X$
		5	11		
			40	22	
		14	28		
			20	10	
		26		30	
			84	48	
		52		76	